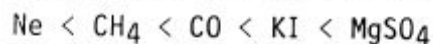


AP Chemistry Intermolecular Forces Supplemental Worksheet Answers

1.

Increasing strength of intermolecular forces



Neon and methane are nonpolar species and possess only induced dipole-induced dipole interactions. Neon has a smaller molar mass than CH_4 , and therefore weaker London (dispersion) forces. Carbon monoxide is a polar molecule. These molecules would be attracted to each other by dipole-dipole interactions. Both KI and MgSO_4 are ionic compounds and the charges associated with the magnesium and sulfate ion (2^+ , 2^-) create greater interparticle forces than those of the potassium and iodide ions (1^+ , 1^-). The lower molecular weight molecules with weaker interparticle forces should be gases at 25°C at 1 atmosphere: Ne , CH_4 , CO .

2.

	<u>Higher Boiling Substance</u>	<u>Intermolecular Force</u>
a.	O_2	induced dipole-induced dipole (London) (greater for heavier molecules)
b.	SO_2	dipole-dipole (stronger than London)
c.	HF	hydrogen bonding (stronger than dipole-dipole)
d.	GeH_4	induced dipole-induced dipole (greater for larger molecules)

3.

	<u>Molecules</u>	<u>Intermolecular Force</u>
a.	CO_2	induced dipole-induced dipole (London)
b.	NH_3	dipole-dipole (Hydrogen bonding)
c.	CHCl_3	dipole-dipole
d.	CCl_4	induced dipole-induced dipole (London)

- 4.
- increase -- The greater the intermolecular forces, the greater the energy (temperature) necessary to "separate" the molecules.
 - decrease -- The stronger interparticle forces will reduce the number of liquid molecules escaping into the gaseous state -- reducing the vapor pressure.
 - not change -- Pressure is force per unit area. If you increase the surface area of the liquid, you also increase the force so the pressure does not change.

5.

$$1.00 \text{ kg CCl}_3\text{F} \cdot \frac{1.00 \times 10^3 \text{ g CCl}_3\text{F}}{1.00 \text{ kg CCl}_3\text{F}} \cdot \frac{1 \text{ mol CCl}_3\text{F}}{137.4 \text{ g CCl}_3\text{F}} \cdot \frac{24.8 \text{ kJ}}{1 \text{ mol CCl}_3\text{F}} = 180. \text{ kJ}$$

6.

- The vapor pressure of water at 60 °C from Figure 11.15 is approximately 150 mm Hg. From Appendix F, the vapor pressure is 149.4 mm Hg.
- At 95 °C the equilibrium vapor pressure of water will be 600 mm Hg.
- The vapor pressure for water at 70 °C is approximately 220 mm Hg and for ethyl alcohol at 70 °C is approximately 520 mm Hg.
- At 20 °C the vapor pressure of diethyl ether is approximately 400 mm Hg.

7.

The vapor pressure of C_2H_5O at 30° is 632 mm Hg.

Calculate the amount of C_2H_5O to furnish this vapor pressure at $30^\circ C$.

$$n = \frac{PV}{RT} = \frac{632 \text{ mm} \cdot \frac{1 \text{ atm}}{760 \text{ mm}} \cdot 0.1 \text{ L}}{0.082 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \cdot 303 \text{ K}} = 3.3 \times 10^{-3} \text{ mol}$$

The total mass of C_2H_5O present (1.00 g) corresponds to 0.014 mol.

The amount of C_2H_5O in the liquid state is $(0.014 - 0.0033)$ 0.0106 mol or 0.48 g. So at 30° approximately 0.52 g is in the gaseous state.

As the flask is cooled from $30^\circ C$ to $0^\circ C$ some of the **gaseous ether will condense** to form liquid ether.

The gas pressure at $0^\circ C$ may be calculated with the Clausius-Clapeyron equation.

$$\ln \frac{P_2}{P_1} = \frac{\Delta H_{\text{vaporization}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\ln \frac{P_2}{632} = \frac{30.95 \text{ kJ/mol}}{8.3143 \times 10^{-3} \text{ kJ/mol} \cdot K} \left(\frac{1}{303} - \frac{1}{273} \right)$$

$$P_2 = 164 \text{ mm Hg}$$

8.

$$\begin{aligned} \text{Volume of room} &= 3.0 \times 10^2 \text{ cm} \cdot 2.5 \times 10^2 \text{ cm} \cdot 2.5 \times 10^2 \text{ cm} \\ &= 1.9 \times 10^7 \text{ cm}^3 \end{aligned}$$

Convert volume to L

$$1.9 \times 10^7 \text{ cm}^3 \cdot \frac{1 \text{ L}}{1.0 \times 10^3 \text{ cm}^3} = 1.9 \times 10^4 \text{ L}$$

To produce a pressure of 60 mm Hg, calculate the amount of ethanol required.

$$P = 60 \text{ mm} \cdot \frac{1 \text{ atm}}{760 \text{ mm}} = 7.9 \times 10^{-2} \text{ atm}$$

$$V = 1.9 \times 10^4 \text{ L}$$

$$T = 25 + 273 = 298 \text{ K}$$

$$\frac{PV}{RT} = \frac{(7.9 \times 10^{-2} \text{ atm})(1.9 \times 10^4 \text{ L})}{(0.082 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}})(298 \text{ K})} = 6.1 \times 10^1 \text{ mol C}_2\text{H}_5\text{OH}$$

$$6.1 \times 10^1 \text{ mol C}_2\text{H}_5\text{OH} \cdot \frac{46 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 2.8 \times 10^3 \text{ g C}_2\text{H}_5\text{OH}$$

This mass of ethanol would occupy a volume of

$$2.8 \times 10^3 \text{ g C}_2\text{H}_5\text{OH} \cdot \frac{1 \text{ cm}^3}{0.79 \text{ g}} = 3.6 \times 10^3 \text{ cm}^3$$

As only 1 L of $\text{C}_2\text{H}_5\text{OH}$ ($1.0 \times 10^3 \text{ cm}^3$) was introduced into the room, **all the ethanol would evaporate.**

9.

The ionic radii for Rb^+ and Cl^- are reported as 148 pm and 181 pm, respectively. The unit cell would have alternating rubidium and chloride ions. The edgest of the unit cell would contain a net of one of each type ion. Their diameters would give a length of 658 pm. The volume of this cell would be

$$V = 2 \times 2 \times 2 = (658 \text{ pm})^3 = 2.85 \times 10^8 \text{ pm}^3 \\ = 2.85 \times 10^{-22} \text{ cm}^3$$

One unit cell would contain 4 ion-pairs. The mass of these ion-pairs is:

$$\frac{120.9 \text{ g RbCl}}{1 \text{ mol RbCl}} \cdot \frac{1 \text{ mol RbCl}}{6.023 \times 10^{23} \text{ RbCl ion pairs}} \cdot 4 \text{ ion pairs} \\ = 8.031 \times 10^{-22} \text{ g RbCl}$$

The density would be

$$D = \frac{M}{V} = \frac{8.031 \times 10^{-22} \text{ g RbCl}}{2.85 \times 10^{-22} \text{ cm}^3} = 2.82 \text{ g/cm}^3$$

10.

$$\text{Solid's density} = 0.77 \text{ g/cm}^3$$

$$\text{Edge of unit cell} = 4.086 \times 10^{-8} \text{ cm}$$

$$\text{Volume of unit cell} = (4.086 \times 10^{-8} \text{ cm})^3 = 6.822 \times 10^{-23} \text{ cm}^3$$

$$\text{Mass} = \text{Density} \cdot \text{Volume}$$

$$= 0.77 \text{ g/cm}^3 \cdot 6.822 \times 10^{-23} \text{ cm}^3 = 5.253 \times 10^{-23} \text{ g}$$

1 LiH ion pair has a mass of

$$\frac{7.94 \text{ LiH}}{1 \text{ mol LiH}} \cdot \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ ion pairs}} = 1.32 \times 10^{-23} \text{ g}$$

$$5.253 \times 10^{-23} \text{ g/unit cell} \cdot \frac{1 \text{ LiH ion pair}}{1.32 \times 10^{-23} \text{ g}} = 3.98 \text{ or } 4 \text{ ion pairs/unit cell}$$

The face-centered cubic lattice is appropriate for LiH.

11.

To melt a 15.0 g ice cube:

$$15.0 \text{ g water} \cdot \frac{1.00 \text{ mol water}}{18.0 \text{ g water}} \cdot \frac{6.01 \text{ kJ}}{1 \text{ mol water}} = 5.01 \text{ kJ}$$

12.

a. Heating 10. kg of liquid ammonia from -50.0°C to -33.6°C requires:

$$10. \times 10^3 \text{ g NH}_3 \cdot 4.7 \text{ J/g}\cdot\text{K} \cdot 16.4 \text{ K} = 7.71 \times 10^5 \text{ J}$$

b. Converting the liquid ammonia to gaseous ammonia requires:

$$10. \times 10^3 \text{ g NH}_3 \cdot \frac{1 \text{ mol NH}_3}{17.0 \text{ g NH}_3} \cdot \frac{25.1 \text{ kJ}}{1 \text{ mol NH}_3} \cdot \frac{1.0 \times 10^3 \text{ J}}{1 \text{ kJ}} \\ = 1.48 \times 10^7 \text{ J}$$

c. Heating the NH_3 vapor from -33.6°C to 0°C requires:

$$10. \times 10^3 \text{ g NH}_3 \cdot 2.2 \text{ J/g}\cdot\text{K} \cdot 33.6 \text{ K} = 7.39 \times 10^5 \text{ J}$$

$$\text{Total energy input} = 7.71 \times 10^5 \text{ J} + 1.48 \times 10^7 \text{ J} + 7.39 \times 10^5 \text{ J} \\ = 1.6 \times 10^7 \text{ J}$$

13.

- At 1.0 atm pressure and room temperature xenon is in the **gaseous phase**.
- At 0.75 atm pressure and -114 °C xenon is in the **liquid phase**.
- The vapor pressure of 380 mm Hg (0.5 atm) is found at a temperature of approximately **-117 °C**.
- At -122 °C the vapor pressure of the solid is approximately **0.15 atm**.
- The **solid phase of xenon is more dense** than the liquid phase. The liquid-solid interface has a **positive slope**. Remember the liquid-solid interface for water has a negative slope-- and that liquid water is more dense than solid water.

14.

$$a. \quad 190 \text{ g C}_4\text{H}_{10} \cdot \frac{1 \text{ mol C}_4\text{H}_{10}}{58.0 \text{ g C}_4\text{H}_{10}} \cdot \frac{24.3 \text{ kJ}}{1 \text{ mol C}_4\text{H}_{10}} = 79.6 \text{ kJ}$$

- Vapor pressure at 0 °F (-17.8 °C or 255.2 K)

From Table 11.5, ΔH vaporization = 24.3 kJ/mol
and vapor pressure = 760 mm at -0.5 °C

$$\ln \frac{P_2}{P_1} = \frac{\Delta H}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

Substituting $T_1 = 272.5 \text{ K}$ $P_1 = 760 \text{ mm Hg}$

$T_2 = 255.2 \text{ K}$ $P_2 = X$

$$\ln \frac{P_2}{760} = \frac{24.3 \text{ kJ/mol}}{8.3143 \times 10^{-3} \text{ kJ/K}\cdot\text{mol}} \left(\frac{1}{272.5} - \frac{1}{255.2} \right)$$

$$\ln \frac{P_2}{760} = -0.727$$

$$\frac{P_2}{760} = 0.483$$

$$P_2 = 367 \text{ mm Hg}$$

The vapor pressure is very low at 0 °F. This means that little C_4H_{10} vaporizes at this temperature, and consequently the camp stove will work poorly.